

17.27 C_6H_6 freezes at 5.5°C . At what temperature will a solution of 10.0 g of

C_4H_{10} in 200 g of C_6H_6 freeze?

The

$$(10.0 \text{ g C}_4\text{H}_{10}) \left(\frac{1 \text{ mol C}_4\text{H}_{10}}{58.0 \text{ g C}_4\text{H}_{10}} \right) = 0.172 \text{ mol C}_4\text{H}_{10} \quad \frac{0.172 \text{ mol C}_4\text{H}_{10}}{0.200 \text{ kg C}_6\text{H}_6} = 0.860 \text{ m}$$

$$\Delta t = (5.12^\circ\text{C/m})(0.860 \text{ m}) = 4.40^\circ\text{C}$$

$$t = (5.5 - 4.4)^\circ\text{C} = 1.1^\circ\text{C}$$

molal

freezing point depression constant of C_6H_6 is 5.12°C/m .

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$$m = \Delta t_b / K_b = \Delta t_f / K_f = (0.50^\circ\text{C}) / (0.51^\circ\text{C/m}) = \Delta t_f / (1.86^\circ\text{C/m})$$

$$\Delta t_f = 1.82^\circ \quad \text{from which } t = -1.82^\circ\text{C}$$

17.28 An aqueous solution

boils

at 100.50°C .

What is the freezing point of the solution?

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17.29 Explain by use of a phase diagram why the normal boiling point of water is raised by addition of a nonvolatile solute while the freezing point is lowered. Label all pertinent point explicitly.

- See Fig. 17.2 the vapor pressure is lowered so that the temperature at which it reaches 1 atm is higher (point A). The intersection of the curve with the solid-vapor curve (which is not affected by solute) is at lower temperature (point B).

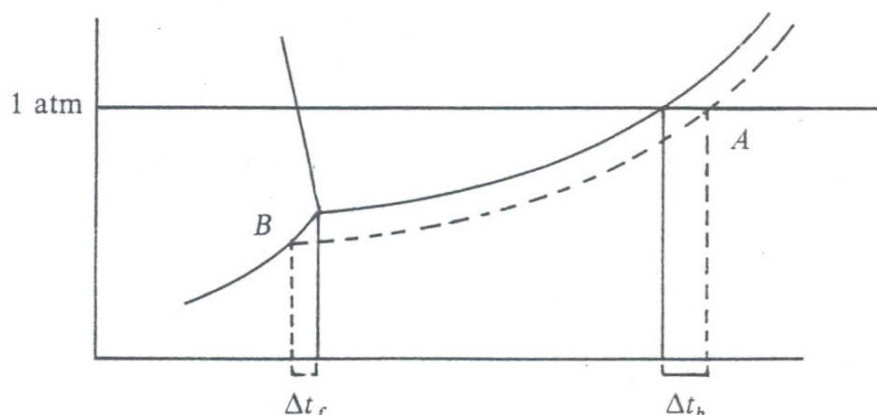


Fig. 17.2

17.30 The freezing point of pure camphor is 178.4°C , and its molal freezing point constant, K_f , is $40.0^{\circ}\text{C} \cdot \text{kg}/\text{mol}$. Find the freezing point of a solution

The first step is to find the molality (m) of the solution.

$$m = \frac{\text{mol of solute}}{\text{kg of solvent}} = \frac{(1.50/125) \text{ mol solute}}{(35/1000) \text{ kg solvent}} = 0.343 \text{ mol/kg}$$

$$\text{Lowering of freezing point} = \Delta t_f = K_f m = (40.0^{\circ}\text{C} \cdot \text{kg/mol})(0.343 \text{ mol/kg}) = 13.7^{\circ}\text{C}$$

$$\text{Freezing point of solution} = (\text{freezing point of pure solvent}) - \Delta t_f = (178.4 - 13.7)^{\circ}\text{C} = 164.7^{\circ}\text{C}$$

containing 1.50 g of a compound of molecular weight 125 in 35.0 g of camphor.

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17.31 Bromoform has a normal freezing point of 7.80°C and its $K_f = 14.4^{\circ}\text{C}/\text{m}$.

$$\text{A} \quad \Delta t = (7.80 - 5.43)^{\circ}\text{C} = 2.37^{\circ}\text{C} \quad m = \frac{\Delta t}{K_f} = \frac{2.37^{\circ}\text{C}}{14.4^{\circ}\text{C}/\text{m}} = 0.165 \text{ m} \quad \frac{2.58 \text{ g}/0.100 \text{ kg}}{0.165 \text{ mol/kg}} = 156 \text{ g/mol}$$

solution of 2.58 g of an unknown in 100 g of bromoform freezes at 5.43°C .

What is the molecular weight of the unknown?

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17.32 When 30.0 g of a nonvolatile solute having the empirical formula CH_2O is dissolved in 800 g of water, the solution freezes at -1.16°C . What is the molecular formula of the solute? $K_f = 1.86^\circ\text{C}/\text{m}$.

- The molality is given by

$$m = \frac{\Delta t}{K_f} = \frac{1.16^\circ\text{C}}{1.86^\circ\text{C}/\text{m}} = 0.624 \text{ m} \quad \frac{30.0 \text{ g}}{0.800 \text{ kg solvent}} = \frac{37.5 \text{ g}}{\text{kg solvent}}$$

Hence 37.5 g is equivalent to 0.624 mol.

$$\frac{37.5 \text{ g}}{0.624 \text{ mol}} = 60.1 \text{ g/mol}$$

The empirical formula weight is 30 g/eq. formula unit. There must be two units per molecule; the formula is $\text{C}_2\text{H}_4\text{O}_2$.

17.33 The freezing point of a solution containing 2.40 g of a compound in 60.0 g of benzene is 0.10°C lower than that of pure benzene. What is the

$$m = \frac{\Delta t}{K_f} = \frac{0.10^\circ\text{C}}{5.12^\circ\text{C}/\text{m}} = 0.0195 \text{ m} \quad \left(\frac{2.40 \text{ g}}{60.0 \text{ g benzene}} \right) \left(\frac{10^3 \text{ g benzene}}{\text{kg benzene}} \right) = 40.0 \text{ g/kg benzene}$$

molecular weight of the compound? (K_f is $5.12^\circ\text{C}/\text{m}$ for benzene.)

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Thus there is 40.0 g of solute and 0.0195 mol of solute per kg of benzene.

$$\frac{40.0 \text{ g}}{0.0195 \text{ mol}} = 2050 \text{ g/mol}$$

17.34 The freezing point of a solution containing 4.80 g of a compound in

60.0

g of

$$\Delta t_f = (5.5 - 4.5) ^\circ\text{C} = 1.0 ^\circ\text{C}$$

$$m = \frac{\Delta t_f}{K_f} = \frac{1.0 ^\circ\text{C}}{5.1 ^\circ\text{C/m}} = 0.20 \text{ m} = 0.20 \text{ mol/kg} \quad M_{\text{compound}} = \frac{4.8 \text{ g/60.0 g benzene}}{0.20 \text{ mol/1000 g benzene}} = 400 \text{ g/mol}$$

benzene is $4.50 ^\circ\text{C}$. What is the molecular weight of the compound?

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17.35 An aqueous solution containing 288 g of a nonvolatile compound having the stoichiometric composition $\text{C}_n\text{H}_{2n}\text{O}_n$ in 90.0 g of water boils at

$101.24 ^\circ\text{C}$

at

$$\Delta t_b = (101.24 - 100.00) ^\circ\text{C} = 1.24 ^\circ\text{C}$$

1.00 atm

$$m = \frac{\Delta t_b}{K_b} = \frac{1.24 ^\circ\text{C}}{0.512 ^\circ\text{C/m}} = 2.42 \text{ m} \quad M = \frac{288 \text{ g/90.0 g water}}{2.42 \text{ mol/1000 g water}} = 1320 \text{ g/mol}$$

pressure.

the

What is

formula of the compound?

molecular

• T

The weight of a molecule is 1320 u. if the value of n were 1, corresponding to CH_2O , the weight would be 30 u. hence n must be $1320/30 = 44$, and the molecular formula is $\text{C}_{44}\text{H}_{88}\text{O}_{44}$.

17.36 Calculate the molecular weight of a substance which forms a 7.0% by mass solution in water which freezes at -0.89

$^{\circ}\text{C}$.

$$\text{Molality} = \frac{\Delta t}{K_f} = \frac{0.89^{\circ}}{1.86^{\circ}/\text{m}} = 0.48 \text{ m}$$

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The solution is 0.48 mol/kg of solvent, and also 70 g/kg of solvent. Thus 70 g is 0.48 mol. The molecular weight is

$$\frac{70 \text{ g}}{0.48 \text{ mol}} = 150 \text{ g/mol}$$

17.37 Calculate (a) the freezing point of a solution of 0.0100 mol of glucose dissolved in 100 g of water, and (b) the freezing point of a 0.100 m solution of naphthalene in benzene.

(a) $\text{Concentration} = \frac{0.0100 \text{ mol}}{0.100 \text{ kg}} = 0.100 \text{ m}$

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$$\Delta t_f = (1.86^{\circ}\text{C/m})(0.100 \text{ m}) = 0.186^{\circ}\text{C}$$

Since the freezing point of pure water is 0.000°C , the freezing point of the solution will be -0.186°C .

(b) $\Delta t_f = (5.12^{\circ}\text{C/m})(0.100 \text{ m}) = 0.512^{\circ}\text{C}$

The freezing point is lowered by 0.512°C , from 5.5 to 5.0°C .

17.38 A solution containing 3.50 g solute X in 50.0 g of water has a volume of 52.5 mL and a freezing point of -0.86°C . (a) Calculate the molality, mole fraction, and molarity of X. (b) calculate the molecular weight of X.

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Thus there is 0.46 mol of X/kg of water, or 0.46 mol of X per 55.5 mol of water. The mole fraction is therefore

$$x(X) = \frac{0.46 \text{ mol}}{(0.46 + 55.5) \text{ mol}} = 0.0082$$

The solution has a density of

$$\frac{(50.0 \text{ g H}_2\text{O}) + (3.50 \text{ g X})}{52.5 \text{ mL}} = 1.02 \text{ g/mL}$$

Per kg of water there is

$$\left(\frac{3.50 \text{ g X}}{50.0 \text{ g H}_2\text{O}} \right) (1000 \text{ g H}_2\text{O}) = 70.0 \text{ g X}$$

The total mass of the solution is thus 1070 g; its total volume is

$$(1070 \text{ g}) \left(\frac{1 \text{ mL}}{1.02 \text{ g}} \right) = 1050 \text{ mL} = 1.05 \text{ L}$$

Then

$$\text{molarity} = \frac{0.46 \text{ mol}}{1.05 \text{ L}} = 0.44 \text{ M}$$

$$(b) \quad \frac{70.0 \text{ g}}{0.46 \text{ mol}} = 152 \text{ g/mol}$$

17.39 Camphor, $\text{C}_{10}\text{H}_{16}\text{O}$, which has a freezing point of 174°C , has a freezing point depression constant of $40.0^\circ\text{C}/\text{m}$. Explain the usefulness and the limitations of camphor as a solvent for determine of molecular weights. For what kind(s) of solutes would camphor be especially useful?

- The freezing point depression constant is so large that an accurate freezing point depression can be obtained on a solution with very few moles dissolved per kilogram of camphor. Thus, high molecular weight materials give measurable freezing point depressions in this solvent. There is a problem in getting many materials to dissolve in camphor, however.

17.40 When 36.0 g of a solute having the empirical formula CH_2O is

dissolved in 1.20 kg of water, the solution freezes at -0.93°C . What is the molecular formula of the solute?

$$m = \frac{\Delta t}{K_f} = \frac{0.93^\circ}{1.86^\circ/\text{m}} = 0.50 \text{ m}$$

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Per kg of solvent, there is 0.50 mol of solute, which has a mass of 36.0 g:

$$\frac{36.0 \text{ g}}{1.20 \text{ kg}} = 30.0 \text{ g/kg} \quad \frac{30.0 \text{ g}}{0.50 \text{ mol}} = 60 \text{ g/mol}$$

The empirical formula weight is 30 g/empirical formula unit; there must be $60/30 = 2$ empirical formula units per molecule. The formula is thus $(\text{CH}_2\text{O})_2$ or $\text{C}_2\text{H}_4\text{O}_2$.

17.41 Calculate the freezing point of 0.200 m solution of fructose, $\text{C}_6\text{H}_{12}\text{O}_6$, (a) in water (b) in acetic acid

- (a) $\Delta t = K_f m = (1.86^\circ\text{C/m})(0.200 \text{ m}) = 0.372^\circ\text{C}$ $t = (0.000 - 0.372)^\circ\text{C} = -0.372^\circ\text{C}$

(b) $\Delta t = K_f m = (3.9^\circ\text{C/m})(0.200 \text{ m}) = 0.78^\circ\text{C}$ $t = (16.6 - 0.78)^\circ\text{C} = 15.8^\circ\text{C}$

17.42 A solution of 10.0 g of a nonionic solute in 100 g of benzene freezes at 4.2°C . Calculate the molecular weight of the solute.

- The freezing point depression, Δt , is $(5.5 - 4.2)^\circ\text{C} = 1.3^\circ\text{C}$

$$m = \frac{\Delta t}{K} = \frac{1.3^\circ\text{C}}{5.12^\circ\text{C/m}} = 0.25 \text{ m}$$

The solution contains 100 g of solute/kg of benzene. It contains 0.25 mol solute/kg of benzene, as determined from the freezing point depression. Thus the molecular weight is

$$\frac{100 \text{ g}}{0.25 \text{ mol}} = 400 \text{ g/mol}$$

17.43 Distinguish (a) between the boiling point of a liquid and the normal boiling point of the liquid, (b) between freezing point and freezing point depression.

- (a) The boiling point is the temperature at which the vapor pressure of the liquid equals the surrounding pressure. The normal boiling point is the temperature at which the vapor pressure is 1 atm. (b) the freezing point is the temperature at which solid first crystallizes from solution; the freezing point depression is the difference between that temperature and the freezing point of the pure solvent. For example, a certain benzene solution starts to freeze at 4.5°C ; its freezing point depression is $(5.5-4.5)^{\circ}\text{C} = 1.0^{\circ}\text{C}$.

17.44 A solution containing 6.35 g of a nonelectrolyte dissolved in 500 g of water freezes at -0.465°C . Determine the molecular weight of the solute.

- $$\Delta t = Km \quad m = \frac{\Delta t}{K} = \frac{0.465^{\circ}\text{C}}{1.86^{\circ}\text{C/m}} = 0.250 \text{ m} = \frac{0.250 \text{ mol}}{\text{kg solvent}}$$

$$\frac{6.35 \text{ g solute}}{500 \text{ g solvent}} = \frac{12.7 \text{ g solute}}{\text{kg solvent}} \quad \text{MW} = \frac{12.7 \text{ g}}{0.250 \text{ mol}} = 50.8 \text{ g/mol}$$

17.45 A solution containing 3.24 g of a nonvolatile nonelectrolyte and 200 g of water boils at 100.130°C at 1 atm. What is the molecular weight of the solute?

- $$\Delta t = Km \quad m = \frac{\Delta t}{K} = \frac{(100.130 - 100.000)^{\circ}\text{C}}{0.513^{\circ}\text{C/m}} = 0.253 \text{ m} = \frac{0.253 \text{ mol}}{\text{kg solvent}}$$

$$\frac{3.24 \text{ g solute}}{200 \text{ g solvent}} = \frac{16.2 \text{ g solute}}{\text{kg solvent}} \quad \frac{16.2 \text{ g}}{0.253 \text{ mol}} = 64.0 \text{ g/mol}$$

17.46 calculate the freezing point and the boiling point at 1 atm of a solution containing 30.0 g cane sugar (molecular weight 342 g/mol) and 150 g

$$(30.0 \text{ g}) \left(\frac{1 \text{ mol}}{342 \text{ g}} \right) = 0.0877 \text{ mol} \quad \frac{0.0877 \text{ mol}}{0.150 \text{ kg}} = 0.585 \text{ m}$$

$$\Delta t_f = K_f m = (1.86 \text{ }^\circ\text{C/m})(0.585 \text{ m}) = 1.09 \text{ }^\circ\text{C} \quad \text{fp} = (0 - 1.09) \text{ }^\circ\text{C} = -1.09 \text{ }^\circ\text{C}$$

$$\Delta t = K_b m = (0.513 \text{ }^\circ\text{C/m})(0.585 \text{ m}) = 0.300 \text{ }^\circ\text{C} \quad \text{bp} = (100.000 + 0.300) \text{ }^\circ\text{C} = 100.300 \text{ }^\circ\text{C}$$

water.

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17.47 A solution was made up by dissolving 3.75 g of a pure hydrocarbon in 95.0 g of acetone. The boiling point of pure acetone was observed to be $55.95 \text{ }^\circ\text{C}$, and of the solution, $56.50 \text{ }^\circ\text{C}$. if the molal boiling point constant of acetone is $1.71 \text{ }^\circ\text{C} \cdot \text{kg/mol}$, what is the approximate molecular weight of the hydrocarbon?

- First method

Compute the molality (m) from the boiling-point equation.

$$\Delta t_b = K_b m \quad (56.50 - 55.95) \text{ }^\circ\text{C} = 1.71 m$$

Solving, $m = 0.322 \text{ mol solute/kg solvent}$. Now find the number of moles of solute in the weighed sample.

$$n(\text{solute}) = \left(0.322 \frac{\text{mol solute}}{\text{kg solvent}} \right) (0.0950 \text{ kg solvent}) = 0.0306 \text{ mol solute}$$

Then

$$\text{MW} = \frac{3.75 \text{ g solute}}{0.0306 \text{ mol solute}} = 123 \text{ g/mol}$$

Second method

For the given concentration,

$$1 \text{ g acetone contains } \frac{3.75}{95.0} \text{ g solute} \quad 1000 \text{ g acetone contains } 1000 \left(\frac{3.75}{95.0} \right) = 39.5 \text{ g solute}$$

Since $1.71 \text{ }^\circ\text{C}$ elevation is produced by 1 mol solute per 1000 g acetone, $0.55 \text{ }^\circ\text{C}$ elevation is produced by

$$\frac{0.55}{1.71} = 0.322 \text{ mol solute}$$

per 1000 g acetone. Thus 0.322 mol solute is equivalent to 39.5 g solute, and

$$\text{MW} = \frac{39.5 \text{ g}}{0.322 \text{ mol}} = 123 \text{ g/mol}$$

17.48 The molecular weight of an organic compound is 58.0 g/mol. Compute the boiling point of a solution containing 24.0 g of the solute and 600 g of water, when the barometric pressure is such that pure water boils at

$$\text{Molality} = m = \frac{n(\text{solute})}{\text{number of kg solvent}} = \frac{(24.0/58.0) \text{ mol solute}}{0.600 \text{ kg solvent}} = 0.690 \text{ mol/kg}$$

$$\text{Elevation of boiling point} = \Delta t_b = K_b m = (0.513 \text{ }^\circ\text{C}\cdot\text{kg/mol})(0.690 \text{ mol/kg}) = 0.354 \text{ }^\circ\text{C}$$

$$\text{Boiling point of solution} = (\text{boiling point of water}) + \Delta t_b = (99.725 + 0.354) \text{ }^\circ\text{C} = 100.079 \text{ }^\circ\text{C}$$

99.725 $^\circ\text{C}$.

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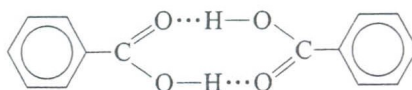
17.49 A certain solution of benzoic acid in benzene has a freezing point 3.1 $^\circ\text{C}$ and a normal boiling point of 82.6 $^\circ\text{C}$. Explain these observations, and

$$m = \frac{\Delta t}{K_f} = \frac{(5.5 - 3.1) \text{ }^\circ\text{C}}{5.12 \text{ }^\circ\text{C/m}} = 0.47 \text{ m} \quad m = \frac{\Delta t}{K_b} = \frac{(82.6 - 80.1) \text{ }^\circ\text{C}}{2.53 \text{ }^\circ\text{C/m}} = 0.99 \text{ m}$$

suggest structure for the solute particles at the two temperatures.

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It is evident that the number of moles of solute particles at the boiling point is twice that at the freezing point of the solution. In fact, benzoic acid dimerizes (forms double molecules) by hydrogen bonding at the lower temperature, but the dimers are broken up at the elevated temperature.



17.50 An aqueous solution of a nonvolatile solute boils at 100.17°C . At what temperature will this solution freeze?

- The molality is calculated from the boiling point elevation:

$$m = \frac{\Delta t}{K_b} = \frac{0.17^{\circ}\text{C}}{0.512^{\circ}\text{C}} = 0.33 \text{ m}$$

The freezing point depression is calculated from that molality:

$$\Delta t = K_f m = (1.86^{\circ}\text{C/m})(0.33 \text{ m}) = 0.62^{\circ}\text{C}$$

The freezing point is

$$t = (0.00 - 0.62)^{\circ}\text{C} = -0.62^{\circ}\text{C}$$

17.51 A certain nonvolatile nonelectrolyte contains 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen. An aqueous solution containing 5.00% by

$$\begin{aligned} (40.0 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.0 \text{ g C}} \right) &= 3.33 \text{ mol C} & (6.7 \text{ g H}) \left(\frac{1 \text{ mol H}}{1.0 \text{ g H}} \right) &= 6.7 \text{ mol H} \\ (53.3 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.0 \text{ g O}} \right) &= 3.33 \text{ mol O} \end{aligned}$$

mass of the
solute boils at
 100.15°C .

Determine the

molecular formula of the compound.

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The empirical formula is CH_2O , which has a formula weight of 30 g/empirical formula unit

$$m = \frac{\Delta t}{K} = \frac{0.15^\circ\text{C}}{0.512^\circ\text{C/m}} = 0.29 \text{ m}$$

$$\frac{50.0 \text{ g}/0.950 \text{ kg H}_2\text{O}}{0.29 \text{ mol/kg H}_2\text{O}} = 180 \text{ g/mol} \quad (2 \text{ significant figures})$$

$$\frac{180}{30} = 6 \text{ units}$$

The molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$.

17.52 Pure benzene freezes at 5.45°C . A solution containing 7.24 g of

C_2H_2	$\Delta t = (5.45 - 3.55)^\circ\text{C} = 1.90^\circ\text{C}$	Cl_4 in
		115.3
g of	$(7.24 \text{ g C}_2\text{H}_2\text{Cl}_4) \left(\frac{1 \text{ mol}}{168 \text{ g}} \right) = 0.0431 \text{ mol}$	
	$\frac{0.0431 \text{ mol C}_2\text{H}_2\text{Cl}_4}{0.1153 \text{ kg C}_6\text{H}_6} = 0.374 \text{ m}$	$K_f = \frac{\Delta t}{m} = \frac{1.90^\circ\text{C}}{0.374 \text{ m}} = 5.08^\circ\text{C/m}$

benzene was observed to freeze at 3.55°C . What is the molal freezing point constant of benzene?

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17.3 OSMOTIC PRESSURE

17.53 The osmotic pressure of blood is 7.65 atm at 37°C . How much glucose should be used per L for an intravenous injection that is to have the same

$$n = \frac{\pi V}{RT} = \frac{(7.65 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(310 \text{ K})} = 0.301 \text{ mol} \quad (0.301 \text{ mol}) \left(\frac{180 \text{ g C}_6\text{H}_{12}\text{O}_6}{\text{mol}} \right) = 54.2 \text{ g}$$

osmotic pressure as blood?

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17.54 What is the osmotic pressure at 0 °C of an aqueous solution containing 46.0 g of (C₃H₈O₃) per L? glycerin

$$\begin{aligned} \pi V &= nRT \\ (46.0 \text{ g}) \left(\frac{1 \text{ mol}}{92.0 \text{ g}} \right) &= 0.500 \text{ mol} \\ \pi &= \frac{(0.500 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(273 \text{ K})}{1.00 \text{ L}} = 11.2 \text{ atm} \end{aligned}$$

17.55 An aqueous solution of urea had a freezing point of -0.52 °C. Predict the osmotic pressure of the same solution at 37 °C. Assume that the molar concentration and the molality are numerically equal.

- The concentration of the solution is not specified, but the effective molality may be inferred from the freezing point lowering.

$$m = \frac{\Delta t_f}{K_f} = \frac{0.52 \text{ }^\circ\text{C}}{1.86 \text{ }^\circ\text{C} \cdot \text{kg/mol}} = 0.280 \text{ mol/kg}$$

The assumption that the molality and molar concentration are equal is not very bad for dilute aqueous solutions. (The relation found in Prob. 16.91 shows that $M \approx m$ when $d \approx 1 \text{ g/cm}^3$ and $W \ll 1000/m$. Urea has a molecular weight of 60 g/mol.) Then 0.280 mol/L may be used for the molar concentration in the osmotic pressure equation.

$$\pi = MRT = (0.280 \text{ mol/L})(0.082 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(310 \text{ K}) = 7.1 \text{ atm}$$

17.56 The osmotic pressure of a solution of a synthetic polyisobutylene in benzene was determined at 25°C. A sample containing 0.20 g of solute/100cm³ of solution developed a rise 2.4 mm at osmotic equilibrium. The density of the solution was 0.88 g/cm³. What is the molecular weight of the polyisibutylene?

- The osmotic pressure is equal to that of a column of the solution 2.4 mm high.

$$\pi = (\text{height})(\text{density})(g) = (2.4 \times 10^{-3} \text{ m})(0.88 \times 10^3 \text{ kg/m}^3)(9.81 \text{ m/s}^2) = 20.7 \text{ Pa}$$

The molar concentration can now be determined from the osmotic pressure equation.

$$M = \frac{\pi}{RT} = \frac{20.7 \text{ N/m}^2}{(8.314 \text{ J/K} \cdot \text{mol})(298 \text{ K})} = 8.3 \times 10^{-3} \text{ mol/m}^3 = 8.3 \times 10^{-6} \text{ mol/L}$$

The solution contained 0.20 g solute/100 mL solution, or 2.0 g/L, and has been found to contain $8.3 \times 10^{-6} \text{ mol/L}$. Then

$$\text{MW} = \frac{2.0 \text{ g}}{8.3 \times 10^{-6} \text{ mol}} = 2.4 \times 10^5 \text{ g/mol}$$

17.57 What would be the osmotic pressure at 17°C of an aqueous solution

$$\text{Molar concentration} = M = \frac{1.75 \text{ g}/(342 \text{ g/mol})}{0.150 \text{ L}} = 0.0341 \text{ mol/L}$$

$$\text{Osmotic pressure} = \pi = MRT = (0.0341 \text{ mol/L})(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(290 \text{ K}) = 0.81 \text{ atm}$$

containing 1.75 g of sucrose (C₁₂H₂₂O₁₁) per 150 cm³ of solution?

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17.58 Calculate the osmotic pressure of a 0.100 M solution of a

$$\pi = \frac{n}{V} RT = \left(\frac{0.100 \text{ mol}}{\text{L}} \right) \left(\frac{0.0821 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273 \text{ K}) = 2.24 \text{ atm}$$

nonelectrolyte at 0°C.

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17.59 What is the molecular weight, M_A , of a solute, A, if the osmotic

$$\pi V = nRT = \frac{m}{M_A} RT \quad M_A = \frac{mRT}{\pi V} = \frac{(10.0 \text{ g})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(300 \text{ K})}{[(10.0/760) \text{ atm}](1.00 \text{ L})} = 18\,700 \text{ g/mol}$$

pressure of a solution containing 10.0 g/L is 10.0 torr at 27°C?

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17.60 A 250-mL water solution containing 48.0 g of sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, at 300 K is separated from pure water by means of semipermeable membrane. What pressure must be applied above the solution in order to just prevent

$$n = \frac{48.0 \text{ g}}{342 \text{ g/mol}} = 0.140 \text{ mol}$$
$$\pi = \frac{nRT}{V} = \frac{(0.140 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(300 \text{ K})}{0.250 \text{ L}} = 13.8 \text{ atm}$$

osmosis?

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17.61 Calculate the osmotic pressure of an aqueous solution which contains

4.00

C₆H₁₂O₆

$$(4.00 \text{ g}) \left(\frac{1 \text{ mol}}{180 \text{ g}} \right) = 0.0222 \text{ mol C}_6\text{H}_{12}\text{O}_6$$

$$\pi = \frac{nRT}{V} = \frac{(0.0222 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(298 \text{ K})}{0.250 \text{ L}} = 2.17 \text{ atm}$$

g glucose,

in 250 mL of

solution at 25°C.

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17.62 A solution of crab hemocynin, a pigmented protein extracted from crabs, was prepared by dissolving 0.750 g in 125 cm³ of an aqueous

medium. At 4

°C

$$n = \frac{\pi V}{RT} = \frac{[2.6 \text{ mm}/(13.6 \times 760 \text{ mm/atm})](0.125 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(277 \text{ K})} = 1.38 \times 10^{-6} \text{ mol}$$

an osmotic

$$\frac{0.750 \text{ g}}{1.38 \times 10^{-6} \text{ mol}} = 5.4 \times 10^5 \text{ g/mol}$$

pressure rise

of

2.6 mm of the

solution was observed. The solution had a density of 1.00 g/cm³. Determine the molecular weight of the protein.

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